**Section 2d reactivity series**

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| 2.15 understand how metals can be arranged in a reactivity series based on their reactions with:  • water • dilute hydrochloric or sulfuric acid.  2.16 understand how metals can be arranged in a reactivity series based on their displacement  reactions between: metals and metal oxides and metals and aqueous solutions of metal salts.  2.17 know the order of reactivity of these metals: potassium, sodium, lithium, calcium, magnesium,  aluminium, zinc, iron, copper, silver, gold.  2.21 *practical: investigate reactions between dilute hydrochloric and sulfuric acids and metals (e.g.*  *magnesium, zinc and iron)* |

**Reaction of metals with cold water**

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| ***Observations:*** the metals that react with cold water are: all group 1 metals, calcium and magnesium (very slowly). Potassium catches fire. | ***Conclusion:*** only reactive metals react with water and this reaction produces an hydroxide and hydrogen as shown by the symbol equations below - the rate at which hydrogen is released can be used to compare reactivity. |

Write balanced symbol equations for the reactions of (i) sodium and water and (ii) lithium with water, include state symbols.

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| (i) |
| (ii) |

**Reaction of metals with steam**

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| Those metals that do not react with cold water can be tested with steam.  Reactive metals react very vigorously with steam therefore steam should only really be used with those metals that show a low rate of reaction with cold water or no reaction at all.  A possible set up for a reaction with steam is shown in the diagram to the right. |  |

(from <http://cbse.myindialist.com/chemistry-x-metals-and-non-metals-what-happens-when-metals-react-with-water/> on 21/1/11)

**Reaction of metals with dilute hydrochloric acid**

Most metals also react with acids. When they do they produce a salt and hydrogen according to the general equation:

metal + hydrochloric acid ⎯→ metal chloride + hydrogen

Write a word equation and a balanced symbol equation for the reaction of calcium with hydrochloric acid and zinc with hydrochloric acid. Include state symbols.

1. calcium + hydrochloric acid:

2. zinc + sulphuric acid:

**Summary of the reactions of metals with water/steam and acid.**

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| --- | --- | --- | --- |
| **metal** | **reaction with cold water** | **reaction with steam** | **reaction with hydrochloric acid** |
| K | fizzes, violently, moves around, catches fire immediately, lilac flame | reaction with steam is too violent | very dangerously explosive |
| Na | fizzes, violently, moves around, might catch fire | dangerously explosive |
| Li | fizzes, moves around | very fast reaction, can be explosive |
| Ca | fizzes slowly | fizzes fast, hydrogen given off |
| Mg | reacts very slowly | reacts to form hydrogen and magnesium oxide | fizz, hydrogen given off |
| Al | no immediate reaction with cold water | reacts to form hydrogen and aluminium oxide | fizz, hydrogen given off |
| Zn | no immediate reaction with cold water | reacts to form hydrogen and an zinc oxide | fizz, hydrogen given off |
| Fe | no immediate reaction with cold water | reacts to form hydrogen and iron oxide | fizz, hydrogen given off |
| Cu | no reaction with cold water | no reaction | no reaction as it cannot displace hydrogen from an acid |

According to the observations in the table above, potassium is the most reactive metal of the metals included in the table and copper is the least reactive. We also know of course that gold and silver are even less reactive than copper.

To summarize further: the order in which the metals appear below is called the reactivity series

|  |  |  |
| --- | --- | --- |
| **reactive unreactive** | | |
| K Na Li Ca Mg Al Zn Fe H Cu Ag Au | | |
| Reacts with acid to produce salt and hydrogen | | no reaction with acid as metal  cannot displace hydrogen from the acids |
| reacts with cold water to produce metal hydroxide and hydrogen | reacts with steam to produce a metal hydroxide/oxide and hydrogen | no reaction with water as it cannot displace hydrogen |

**Displacement reactions**

When a metal is pure and unreacted it exists in the form of atoms. When a metal reacts with a non-metal, the metal forms positive ions. Therefore metals in solutions of ionic compounds (metal salts) are in the form of positive ions.

A displacement reaction is a reaction in which the atoms of a more reactive metal displace the ions of a less reactive metal from its compound to become atoms again. This is because the atoms of the more reactive metal have much greater tendency to form ions than atoms of the less reactive metal. The more reactive the metal, the easier it forms ions. The less reactive a metal, the greater its tendency to exist as atoms.

The table below shows two examples of displacement reactions.

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| Reaction between magnesium atoms and copper sulphate (copper ions – metal salt) to form magnesium ions and copper atoms. Magnesium forms ions much easier than copper. As the copper ions are removed the solution loses its blue colour. **Magnesium displaces copper.** | Reaction between iron atoms and copper sulphate (copper ions) to form iron ions and copper atoms. Iron forms ions much easier than copper. The copper atoms are being deposited onto the iron nail. **Iron displaces copper.** |
|  |  |
| magnesium + copper sulphate  → magnesium sulphate + copper  Mg (s) + CuSO4 (aq) → MgSO4 (aq) + Cu (s) | iron + copper sulphate  → iron (II) sulphate + copper  Fe (s) + CuSO4 (aq) → FeSO4 (aq) + Cu (s) |

Exercises

1. Predict whether or not the following reactions will take place. For those that will take place write word equations and balanced symbol equations.

a) magnesium + copper (II) oxide

b) iron + aluminium oxide

c) calcium + magnesium oxide

1. Predict whether or not the following reactions will take place. For those that will take place write word equations and balanced symbol equations.

a) magnesium + calcium nitrate

b) iron + copper (II) chloride

c) copper + silver nitrate

1. The table below shows the results obtained from a displacement experiment. Place the metals

in order of increasing reactivity with the least reactive first

most reactive: …………….. ………………….. least reactive: ……………….

|  |  |  |  |
| --- | --- | --- | --- |
|  | salt solutions of metals (e.g. sulphates) | | |
| metal | solution of X | solution of Y | solution of Z |
| X | no reaction | reaction | reaction |
| Y | no reaction | no reaction | reaction |
| Z | no reaction | no reaction | no reaction |

**Reduction and oxidation - redox**

1. **Oxidation and reduction in terms of loss and gain of oxygen**

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| 2.18 know the conditions under which iron rusts  2.19 understand how the rusting of iron may be prevented by:  • barrier methods • galvanising  • sacrificial protection.   * 1. understand the terms: oxidation, reduction, redox, oxidising agent, reducing agent in terms of gain or loss of oxygen and loss or gain of electrons. |

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| Oxidation = a reaction during which a substance gains oxygen /gains mass |

#### Examples of oxidation reactions

1. Heating of copper to form copper oxide, which is black powder, the copper is oxidised and its mass increases as oxygen is added.

2Cu (s) + O2 (g) → 2 CuO (s)

1. Burning of magnesium to form magnesium oxide; the magnesium is oxidised and its mass increased.

2Mg (s) + O2 (g) → 2 MgO (s)

Other examples of oxidation reactions:

* rusting or corrosion of iron.
* combustion of fuels
* respiration

#### Examples of reduction reactions:

Heating of mercury oxide; the mercury oxide is reduced as it loses its oxygen (mass is decreased).

2HgO (s) → 2Hg (l) + O2 (g)

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| *Reduction = a reaction during which a substance loses oxygen (loses mass)* |

*A redox reaction = a reaction during which both an oxidation and reduction occur. Oxygen is lost by one reactant and taken by another reactant.*

Example of redox reaction: Reaction between copper oxide and hydrogen.

**reduction** (oxygen is being lost)

CuO (s) + H2 (g) Cu (s) + H2O (g)

**oxidising reducing**

**agent agent**

**oxidation** (oxygen is gained)

In this reaction:

* the copper oxide has been reduced by the hydrogen which has been oxidised;
* the CuO is the oxidising agent;
* the hydrogen is the reducing agent.

In any redox reaction, the reducing agent, which gains the oxygen, causes the reduction of the other substance and by doing so becomes oxidised. The oxidising agent causes the oxidation of the other substance but itself becomes reduced in the process.

Complete the following passage.

When during a reaction a substance, usually an element, gains oxygen we say that the substance has

been \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ and an \_\_\_\_\_\_\_\_\_\_\_\_\_\_ has taken place.

When during a reaction, a compound loses oxygen we say the compound has been \_\_\_\_\_\_\_\_\_\_\_\_\_\_

and a \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ has taken place.

During a reduction often the oxygen that is lost goes to another substance which then becomes

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ . A reaction during which one substance is oxidised and another substance is

reduced is called a \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ reaction.

A substance that causes another substance to lose oxygen and becomes itself oxidised is called a

\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_.

A substance that loses oxygen to another substance becomes \_\_\_\_\_\_\_\_\_\_\_\_\_ and causes the other

substance to be oxidised; such a substance is called a \_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_\_ .

For each of the redox reactions below state which substance is the oxidised and which is reduced. Also identify the oxidising agent and the reducing agent for each reaction.

(a) aluminium + iron oxide ⎯→ iron + aluminium oxide

(b) magnesium + carbon dioxide ⎯→ magnesium oxide + carbon

(c ) tin oxide + carbon ⎯→ tin + carbon monoxide

(d) NO2 (g) + SO2 (g) ⎯→ NO (g) + SO3 (g)

(e) Fe (s) + CuO (s) ⎯→ FeO (s) + Cu (s)

(f) Fe2O3 (s) + 3CO (g) ⎯→ 2Fe (s) + 3CO2 (g)

**Oxidation and reduction in terms of loss and gain of electrons**

OILRIG = **O**xidation **I**s **L**oss (of electrons) **R**eduction **I**s **G**ain (of electrons)

**Redox reaction = a reaction during which an electron(s) is transferred from one substance to another.**

How do we know if a substance is oxidised or reduced? Check oxidation states.

How?

* The oxidation state of an atom =
* its ionic charge if it has formed an ion; or
* the number of electrons it shares with another atom of a different element in a covalent bond;
* the oxidation state of an element on its own is zero.
* The total of all oxidation states in a molecule is zero.
* if its oxidation state increases a substance has been oxidised
* if its oxidation state decreases, a substance has been reduced;
* if there are no changes in oxidation states then no oxidation or reduction has taken place.

Example of how oxidation states in a reaction are used to identify reducing and oxidising agent:

Reaction between copper oxide and hydrogen.

(copper has gained 2 electrons)

**reduction** (oxidation state of Cu has decreased)

CuO (s) + H2 (g) Cu (s) + H2O (g)

+2 -2 0 0 +1 -2

**oxidising reducing**

**agent agent**

**oxidation** (oxidation state has increased) hydrogen has lost electrons

Hydrogen as the reducing agent has reduced copper. Or copper oxide has oxidized hydrogen and is the oxidizing agent.

For each of the reactions below decide if it is a redox reaction or not; if it is a

redox reaction say which reactant is the reducing agent and which one is the oxidising agent:

(1) O2(g) + 2H2 (g) ⎯→ 2H2O (g)

(2) 2Fe(s) + 3Cl2 (g) ⎯→ 2FeCl3 (s)

(3) MgO(s) + 2HCl (aq) ⎯→ MgCl2 (aq) + H2O (l)

(4) Mg (s) + CuO (s) ⎯→ MgO (s) + Cu (s)

(5) Ca(OH)2 (s) + H2SO4 (aq) ⎯→ CaSO4 (s) + 2H2O (l)

(6) CuSO4.5H2O (s) ⎯→ CuSO4 (s) + 5H2O (l)

(7) Fe2O3 (s) + 3H2 (g) ⎯→ 2Fe (s) + 3H2O (g)

(8) 2NaBr (aq) + Cl2 (aq) ⎯→ 2NaCl (aq) + Br2 (aq)

(9) CaCO3 (s) ⎯→ CaO (s) + CO2 (g)

(10) 2Mg(s) + SO2 (g) ⎯→ 2MgO (s) + S (s)

(11) N2(g) + 3H2 (g) ⎯→ 2NH3 (g)

(12) 2AgNO3 (aq) + MgCl2 (aq) ⎯→ 2AgCl (s) + Mg(NO3 )2 (aq)

(13) CaCO3 (s) + 2HNO3 (aq) ⎯→ CO2 (g) + Ca(NO3)2 (aq) + H2O (l)

(14) C(s) + 2H2SO4 (aq) ⎯→ CO2 (g) + SO2 (s) + 2H2O (l)

The displacement reactions of the halogens that you studied in subsection 2b are also examples of redox reactions as the examples below show**.**

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| Cl2 (aq) + 2NaBr (aq) → Br2 (aq) + 2NaCl (aq)  During this reaction, the chlorine displaces the bromine. The chlorine atoms are reduced (each gains an electron) and bromide ions are oxidized (each lose an electron). This displacement reaction is also a redox reaction:  Half equations: reduction: Cl2 + 2e- → 2Cl- oxidation: 2Br- → Br2 +2e-  Ionic equation: Cl2 (aq) + 2Br- (aq) → Br2 (aq) + 2Cl- (aq) |
| Cl2 (aq) + 2NaI (aq) → I2 (aq) + 2NaCl (aq)  During this reaction, the chlorine displaces iodine. The chlorine atoms are reduced (each gain an electron) and iodide ions are oxidized (each loses an electron).  Half equations: reduction: Cl2 + 2e- → 2Cl- oxidation: 2I- → I2 +2e-  Ionic equation: Cl2 (aq) + 2I- (aq) → I2 (aq) + 2Cl- (aq) |
| Br2 (aq) + 2NaI (aq) → I2 (aq) + 2NaBr (aq)  During this reaction, the bromine displaces the iodine. The chlorine atoms are reduced (each gain an electron) and iodide ions are oxidized (each lose an electron).  Half equations: reduction: Br2 + 2e- → 2Br- oxidation: 2I- → I2 +2e-  Ionic equation: Br2 (aq) + 2I- (aq) → I2 (aq) + 2Br- (aq) |

**Rusting**

Rusting is a reaction between iron, oxygen and water producing a brittle red-brown product which is iron (III) oxide which weakens any iron structure which is why this process is an expensive nuisance.

iron + oxygen + water → hydrated iron (III) oxide (=rust) (Fe2O3.xH2O)

The conditions necessary for rusting can be investigated using simple experiments as shown on the following site: <http://www.bbc.co.uk/scotland/learning/bitesize/standard/chemistry/metals/corrosion_rev1.shtml>

The conditions for rusting are: oxygen and water whilst any electrolyte e.g. sodium chloride speeds up the process.

The prevention of rusting should focus on ensuring all three reactants (iron, oxygen and water) do not come in contact with each other.

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| --- | --- | --- | --- |
| method | how it works | method | examples |
| Painting the iron surface, applying oil and grease, covering in plastic | A barrier between iron and oxygen/water. | Cover the iron. However, when barrier is broken, rusting occurs | Bicycle chains (oil), bicycle brake cables (plastic), bridges (paint) |
| Galvanising (= covering steel with zinc) (still works even if zinc is scratched away at some places) | Zinc is more reactive/better at losing electrons than iron so the zinc is oxidized by the oxygen and water instead of the iron | Zinc can be scratched and iron will still not rust.  Need to keep checking if there is enough zinc left on the iron. | Watering cans |
| Sacrificial protection with a more reactive metal (e.g. zinc, aluminium or magnesium) (zinc is used in ships) (galvanizing is also a form of sacrificial protection) | More reactive metal is oxidized by the oxygen and water instead of the iron. | More reactive metal can be scratched and iron will still not rust.  Need to keep checking if there is enough zinc left on the iron. | Ships, underground pipes |